Ch. 12: Stoichiometry 5: Excess Reactant & Amount of Product Formed Dr. Casagrande

Follow along as you view the video, "Stoichiometry: Excess Reactant & Amount of Product Formed" on edpzzle.com and fill in the blanks as you go. (Also available at (http://youtu.be/jbFGSUi1GLQ)

- Amount of Excess Reactant & Amount of Product
  - Determining amount of excess reactant remaining
    - Use mass-mass calculation starting with limiting reactant (G) to determine amount of excess reactant used (W)
    - Subtract amount used from amount present to find amount remaining
  - Determining amount of product formed
    - Use mass-mass calculation starting with limiting reactant (G) to determine amount of product formed (W)
  - Mole-Based Problem
    - In previous lesson, when 2.3 mol N_2 and 7.6 mol H_2 reacted according to the reaction

      $$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g),$$

      N_2 was the limiting reactant. How much of the excess reactant (H_2) remained? How much NH_3 is produced?
      - Determine mol H_2 used by 2.3 mol N_2 then subtract:

        $$\text{mol } H_2 \text{ used} = 2.3\ \cancel{\text{mol } N_2} \times \frac{3 \text{ mol } H_2}{1\ \cancel{\text{mol } N_2}} = 6.9 \text{ mol } H_2 \text{ used}$$

        $$\text{mol } H_2 \text{ remaining} = 7.6 \text{ mol } H_2 \text{ available} - 6.9 \text{ mol } H_2 \text{ used} = \boxed{0.7 \text{ mol } H_2 \text{ remain}}$$

      - Next determine mol NH_3 formed from 2.3 mol N_2:

        $$\text{mol } NH_3 = 2.3\ \cancel{\text{mol } N_2} \times \frac{2 \text{ mol } NH_3}{1\ \cancel{\text{mol } N_2}} = \boxed{4.6 \text{ mol } NH_3 \text{ formed}}$$

  - Mass-Based Problem
    - In the last lesson, when 20.0 g N_2 and 10.0 g H_2 reacted by the same reaction, N_2 was the limiting reactant. How much H_2 remains? How many grams of NH_3 are formed?
      - Since we already know the moles of N_2, you can start there):

        $$\text{mass } H_2 \text{ used} = \frac{0.714\ \cancel{\text{mol } N_2}}{} \left| \frac{3\ \cancel{\text{mol } H_2}}{1\ \cancel{\text{mol } N_2}} \right| \frac{2.016 \text{ g } H_2}{1\ \cancel{\text{mol } H_2}} = 4.32 \text{ g } H_2 \text{ used}$$

        $$\text{mass } H_2 \text{ remaining} = 10.0 \text{ g } H_2 \text{ present} - 4.32 \text{ g } H_2 \text{ used} = \boxed{5.68 \text{ g } H_2 \text{ remaining}}$$

        $$\text{mass } NH_3 = \frac{0.714\ \cancel{\text{mol } N_2}}{} \left| \frac{2\ \cancel{\text{mol } NH_3}}{1\ \cancel{\text{mol } N_2}} \right| \frac{17.04 \text{ g } NH_3}{1\ \cancel{\text{mol } NH_3}} = \boxed{24.3 \text{ g } NH_3 \text{ formed}}$$

  - Your Turn
    - In the last lesson, when 84.9 g FeS reacted with 64.9 g O_2 by the reaction

      $$4\,FeS(s) + 7\,O_2(g) \rightarrow 2\,Fe_2O_3(s) + 4\,SO_2(g),$$

      FeS was limiting. How many grams of O_2 remain, and how many grams of Fe_2O_3 are formed?

      $$\text{mass } O_2 \text{ used} = 0.966\ \cancel{\text{mol FeS}} \times \frac{7\ \cancel{\text{mol } H_2O}}{4\ \cancel{\text{mol FeS}}} \times \frac{32.00 \text{ g } O_2}{1\ \cancel{\text{mol } O_2}} = 54.1 \text{ g } O_2 \text{ used}$$

      $$\text{mass } O_2 \text{ remaining} = 64.9 \text{ g } O_2 \text{ present} - 54.1 \text{ g } O_2 \text{ used} = \boxed{10.8 \text{ g } O_2 \text{ remaining}}$$

      $$\text{mass } Fe_2O_3 = 0.966\ \cancel{\text{mol FeS}} \times \frac{2\ \cancel{\text{mol } Fe_2O_3}}{4\ \cancel{\text{mol FeS}}} \times \frac{159.7 \text{ g } Fe_2O_3}{1\ \cancel{\text{mol } Fe_2O_3}} = \boxed{77.1 \text{ g } Fe_2O_3 \text{ formed}}$$

  - Read §12.3 pp. 364-369 for additional sample problems